Year 12 ATAR Chemistry

Unit 3: Acid Base Titrations
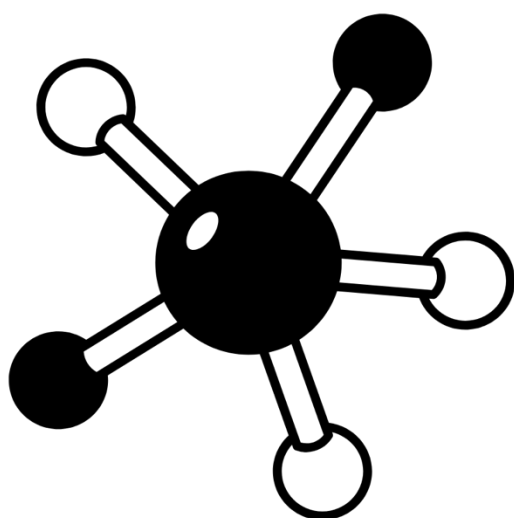
Acids and Bases

Chemistry ATAR Year 12 Unit 3

Acid Base Titrations

Syllabus Points Covered

This unit of work covers volumetric analysis used in acid-base titrations. Data obtained during titration can be used to calculate information about unknown solutions. Volumetric analysis methods involving acid-base reactions rely on the identification of an equivalence point by measuring the associated change in pH, using appropriate acid-base indicators or pH meters, to reveal an observable end point.

Read the information below:

Volumetric Analysis

Acid-base reactions can be used to determine information about unknown solutions. An example could be the concentration of acetic (ethanoic) acid in a sample of vinegar or the acidity of a sample of biodiesel.

In **volumetric analysis**, in order to work out the concentration of the acid in the solution, a chemist will react the acid with an alkaline solution. If the chemist knows the concentration of the alkali, and measures the volume of the alkali used to neutralise all of the acid, the chemist can calculate the number of moles of alkali used. Once this is known, the unknown number of moles of the acid can be worked out. If the volume of this acid is measured, the concentration of the acid can be worked out.

Primary Standard Solutions

Acid base titrations involve a **controlled neutralisation** reaction between an acid solution and a base solution.

The **concentration** of either the acid solution or the base solution is **known** with great accuracy, (to at least three significant figures). This is the **standard solution** (primary or secondary).

The concentration of the other solution is **unknown**. The purpose of the titration is to determine the concentration of the acid or base in the unknown solution.

To ensure that a primary standard solution is prepared with great accuracy it must use a suitable primary standard. There a number of features that are important for a primary standard:

- obtained with a high degree of purity and has a known formula

- undergoes reactions according to known chemical reactions

- Must be stable. Composition does not change on exposure to atmosphere. These substances are not a primary standard: NaOH **(deliquescence)** and $Na_2CO_3.10H_2O$ **(hygroscopic).**

- should have a relatively high formula mass to minimise weighing errors.

NOTE: a primary standard solution and a primary standard are not the same.

A standard solution can be primary or secondary.

- A **primary standard solution** is prepared by accurately weighing out a **primary standard**, then making this up to a required concentration using distilled water.

- A **secondary standard solution** can be prepared by titration against a primary standard solution.

To make a **primary standard solution** there a number of steps to follow:

- accurately weigh out a previously calculated mass of a primary standard

- quantitatively transfer the solid to a volumetric flask

- add a small amount of distilled water and swirl to dissolve the solid

- add more distilled water to make the solution up to the calibration mark on the neck of the flask (bottom of meniscus on the mark).

- invert the flask a few times to make sure the concentration is even throughout

Finding the Concentration of a Primary Standard Solution

Example: A sodium carbonate primary standard solution is prepared by dissolving 3.747 g of Na_2CO_3 solid in approximately 100 mL of water and then making up to the mark in a 500 mL volumetric flask. Determine the concentration of this solution.

Na_2CO_3 is heated in an oven over 200 °C to remove moisture.	Weigh out Na_2CO_3.	Transfer Na_2CO_3 to volumetric flask.	Ensure complete transfer by washing with water.	Dissolve Na_2CO_3 in water.	Add water up to calibrated mark.

Photos by SIDE

Find the
moles of
Na_2CO_3 used.

$n(Na_2CO_3) = \underline{m} = \dfrac{3.747}{[(22.99 \times 2) + 12.01 + (16 \times 3)]} = 0.03535$ mol

Find conc. of
Na_2CO_3

$c(Na_2CO_3) = \underline{n} = \dfrac{0.03535}{0.500} = 7.07 \times 10^{-2}$ mol L^{-1}

Acid base primary standards include

- anhydrous sodium carbonate, Na_2CO_3 (a base) and
- oxalic acid dihydrate, $H_2C_2O_4.2H_2O$ (an acid).

Exercise 1: Standard Solutions

Below are some questions to assist your understanding. Complete Exercise 1 and
check your answers at the end of the book.

a) 1.336 g of pure anhydrous Na_2CO_3 (primary standard) was dissolved in water and
    made up to 250.0 mL of solution in a volumetric flask (primary standard solution).
    Calculate the concentration of this standard solution.

b) 500.0 mL of a 0.250 mol L^{-1} primary standard solution of oxalic acid ($H_2C_2O_4.2H_2O$)
    is required to be made. What mass of primary standard is required?

c) A 5.00 mL sample of commercial vinegar was completely neutralised by 24.50 mL
    of 0.150 mol L^{-1} NaOH solution. Calculate the ethanoic acid concentration of the
    vinegar.

Performing the Acid-Base Titrations

A *titration* is the process in which a solution is added from a <u>burette</u> to another solution
with which it reacts, so that the volume of the added solution can be accurately measured.

One solution is added to another until the reaction between them is complete.

Pipette

A definite volume of one of the solutions is placed in a conical flask using a <u>pipette</u> (e.g.
20.00 mL).

A pipette should be cleaned with soapy water, rinsed with water, rinsed with distilled water
and finally, it should ALWAYS be rinsed with the reagent to be measured. This ensures
that the concentration of the reagent in the pipette is not diluted.

Sketched by SB

Burette

The other solution is added from a <u>burette</u> (in a controlled manner) until the reaction is complete and a colour change is observed, signalling the **end point** of the titration.

A burette is used to deliver a precise volume (and variable) of solution (e.g. 22.35 mL).

A burette should be cleaned in a similar way to a pipette.

Photo by SIDE

Performing the Titration

The six steps used in volumetric analysis to calculate the concentration of a solution are shown below.

Step 1. Start with a solution with a known concentration. (This is called a standard solution.) Accurately measure a sample (e.g. 20.0 mL) using a pipette. Transfer to a cleaned conical flask. Add an appropriate indicator.

Step 2. React the standard solution with a solution that has an unknown concentration. This is done by adding the solution from a burette until a colour change occurs.

Step 3. Record the volumes of both solutions required for a complete reaction.

Step 4. Calculate the number of moles of the standard solution used. (Number of moles equals concentration multiplied by volume in litres ($n = c \times V$).)

Step 5. Use the equation to calculate the number of moles of the unknown solution. (Use the conversion factor 'ratio' from the balanced chemical equation.)

Step 6. Use the measured volume to calculate the unknown concentration of the solution. (Concentration equals the number of moles divided by the volume used in the titration ($c = n / V$).)

Example:

During a titration, 25.0 mL of 0.150 mol L^{-1} sodium hydroxide was measured into a conical flask, using a pipette. 3 drops of indicator were added. Then, using a burette, sulphuric acid of concentration 0.200 mol L^{-1} was added until neutralisation occurred. What volume of the acid was needed?

(a) Write a balance equation

$$H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$$
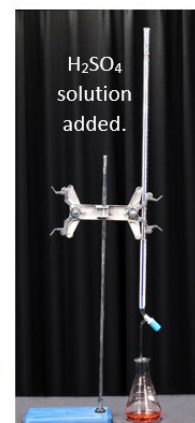
(b) Calculate the number of moles of the known sample – NaOH

n(NaOH) = cV
= 0.150 x 0.0250   = 0.003750 mol

(c) Calculate the number of moles of H$^+$

1 mol H$^+$ reacts with 1 mol of OH$^-$
n(H$^+$) = n(OH$^-$)   = 0.003750 mol H$^+$

(d) Calculate the number of moles of H$_2$SO$_4$

1 mol of H$_2$SO$_4$ has 2 mol H$^+$
n(H$_2$SO$_4$) = ½ n(H$^+$)
= ½ x 0.003750   = 0.001875 mol

(e) Calculate the volume of H$_2$SO$_4$

$$c(H_2SO_4) = \frac{n}{V} = \frac{0.001875}{0.2} = 0.00938\ L = 9.38\ mL$$

Acid-base indicators are weak acids, or weak bases, in which the acidic form is a different colour from the basic form.

Indicators in Acid-Base Titrations

Acid-base indicators are weak acids, or weak bases. They can dissociate in water. When the indicator is in a molecular form it will absorb light to create a particular colour in solution. In the ion form it will create a different colour.
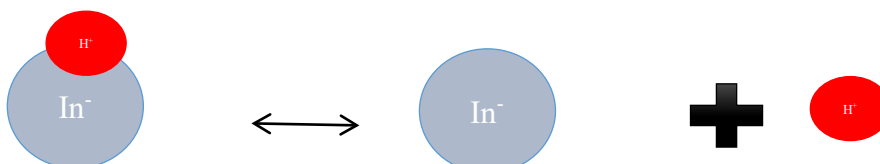
Indicator: Colour A in acid   Indicator ion: Colour B in base   Sketched by SB

An indicator is a substance that changes colour over a specific pH range. In the previous example, in an acidic environment, the indicator will be present mostly in the molecular form. In a basic environment, the H$^+$ ion is removed from the molecule and the ionic form is dominant.

**The colours of indicators in acidic and basic solutions**

Indicator	Colour on acid side	pH at colour change	Colour on basic side
methyl orange	red	3–5	yellow
litmus	red	5–8	blue
phenolphthalein	colourless	8–10	pink

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pH changes

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Consider a 20.0 mL of 1.00 mol L^{-1} ammonia solution. It will have a pH of ~9.

As a small amount of 1.00 mol L^{-1} HCl solution is added, the H$^+$ will neutralise some of the OH$^-$. Since there is a large amount of OH$^-$, the pH changes by only a very small amount.
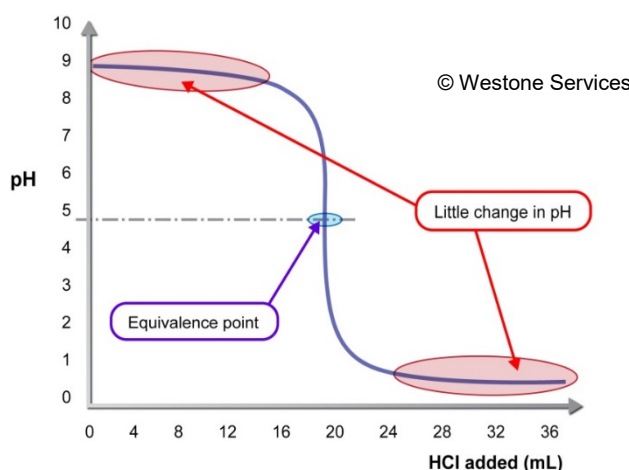
After about 20mL of HCl is added, the pH changes very rapidly – the equivalence point. In this example, the equivalence point is not pH = 7 because the final solution will be slightly acidic.

Beyond this point, there is an excess of H$^+$ and any added HCl will not change the pH by much.

Equivalence Point

Equivalence point – the point at which chemically equivalent amounts of acid and base are added (when acid and base have been added in sufficient volumes so reaction is complete). Both acid and base are fully consumed.

The pH of the titration mixture suddenly changes at this point:

- when a strong acid is titrated with a strong base the pH at the equivalence point is 7
- one drop of NaOH solution can change the pH from 4 to 10
- at this stage the indicator changes colour i.e. the end point is reached.

On a titration curve, the equivalence point is a point roughly half way down the almost vertical part of the curve.
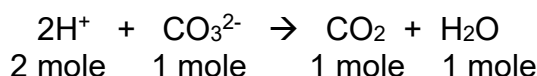
The ratio of moles of acid and base that have been added is the same as the mole ratio in the balanced equation:

e.g. when HCl neutralises NaOH.

$$H^+ \ + \ OH^- \ \rightarrow \ H_2O$$
1 mole   1 mole   1 mole
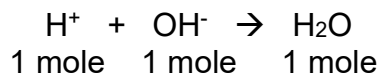
The equivalence point occurs when the number of moles of hydrochloric acid equals the same number of moles of sodium hydroxide.

e.g. when HCl neutralises Na_2CO_3

$$2H^+ \ + \ CO_3^{2-} \ \rightarrow \ CO_2 \ + \ H_2O$$
2 mole   1 mole   1 mole   1 mole

The equivalence point occurs when the number of moles of hydrochloric acid equals twice the number of moles of sodium carbonate.

e.g. when H_2SO_4 neutralises NaOH

$$H_2SO_{4(aq)} \ + \ 2OH^- \ \rightarrow \ SO_4^{2-} \ + \ 2H_2O$$
1 mole   2 mole   1 mole   2 mole
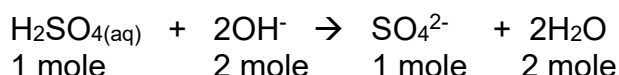
The equivalence point occurs when the number of moles of sodium hydroxide equals twice the number of moles of sulphuric acid.

End Point

This occurs when the indicator **changes colour** during a titration – shows reaction is complete (equivalence point) – if the correct indicator has been selected.

pH at the equivalence point can be acidic or basic because of the reaction products that have been formed.

If the reaction products are acidic (eg. CO_2, NH_4^+) then the end point will be acidic.

If the products are basic (eg. CH_3COO^-, PO_4^{-3}) the end point will be basic.

The indicator chosen should change colour at or near the equivalence point pH.

General Rules:

- Strong acid + strong base – neutral end point
- Strong acid + weak base – acidic end point
- Weak acid + strong base – basic end point
- Weak acid + weak base – neutral end point

Exercise 2: Acid Base Reactions.

Below are some questions to assist your understanding. Complete Exercise 2 and check your answers at the end of the book.

The processes below are all neutralisation reactions between solutions. Write equations for the following reactions.

1. potassium hydroxide + hydrochloric acid → potassium chloride + water

2. ammonia + hydrochloric acid → ammonium chloride

3. sodium hydroxide + acetic acid → sodium acetate + water

4. sodium carbonate + sulfuric acid → sodium sulphate + water + carbon dioxide

Use the products of the reaction to work out the pH of the final solution.

Indicator

Choice of *indicator* depends on;
- pH of the equivalence point (determined by the products remaining in solution)

- range of colour change of indicator.

The indicator must change colour close to the equivalence point and preferably have two distinct colours at different pH.

Strong acid – strong base: equivalence point at pH = 7, neutral.

Weak acid – strong base: at the equivalence point the pH is on the basic side of 7 (i.e. > 7) due to the formation of $OH^-_{(aq)}$ by hydrolysis.

Strong acid – weak base: at the equivalence point the pH is on the acidic side of 7 (i.e. < 7), again by hydrolysis, but this time $H^+_{(aq)}$ or $H_3O^+_{(aq)}$ are formed.

Weak acid – weak base: depends, but equivalence point should be about 7, neutral.

General rule
- s.a. + s.b. = neutral
- s.a. + w.b. = acidic
- s.b. + w.a. = basic
- w.a. + w.b. ~ neutral

Watch the video   https://www.youtube.com/watch?v=-1nJv0k8zQU

Ideally, the end point and the equivalence point occur at the same pH. In practice this may not occur as the indicator colour change rarely coincides precisely with the pH at equivalence.

Type of titration	pH at equiv. point	Indicator	Range of colour change of indicator	Colour
strong acid – strong base e.g. HCl – NaOH	7	bromothymol blue methyl orange phenolphthalein litmus methyl red	3.0 – 4.6 3.1 – 4.4 8.3 – 10.0 5.0 – 8.0 4.4 – 6.2	yellow → blue red → orange colourless →pink red → blue red → yellow
strong acid – weak base e.g. HCl – NH$_3$	~5	methyl orange	3.1 – 4.4	red → yellow
weak acid – strong base e.g. CH$_3$COOH – NaOH	~9	phenolphthalein	8.3 – 10.0	colourless →pink
Weak acid – weak base e.g. CH$_3$COOH – NH$_3$	~7	Phenolphthalein Alizarin yellow	8.3 – 10.0 10.1 – 12.0	colourless →pink yellow→orange

Selection of Indicator
- *Acid Base strength*   **strong acid + strong base**   HCl + NaOH

Explanation of equivalence pH   neutral - need change in neutral region

Suggested indicator   phenolphthalein – pH change at end point with the addition of one more drop - 7 to 10

Colour change when acid is in the flask   colourless to pink

- *Acid Base strength*   **strong acid + weak base**   HCl + NH$_3$

Explanation of equivalence pH   acidic due to NH$_4^+$ - need change in acid region

Suggested indicator   methyl orange

Colour change when acid is in the flask   yellow to red

- *Acid Base strength*   **weak acid + strong base**   CH$_3$COOH + NaOH

Explanation of equivalence pH   Basic due to CH$_3$COO$^-$ ions - change in basic region

Suggested indicator	phenolphthalein
Colour change when acid is in the flask	colourless to pink
• *Acid Base strength* NH$_3$	**weak acid + weak base** CH$_3$COOH +
Explanation of equivalence pH	neutral - need change in neutral region
Suggested indicator	phenolphthalein (difficult to determine)
Colour change when acid is in the flask	colourless to pink

Remember:

End point *= indicator changes colour.*   ***Equivalence point*** *= stoichiometric ratio*

Points to note;

- indicators are selected so that their end point is as close as possible to the reaction's equivalence point

- in a reaction between a strong acid and a strong base the addition of a very small amount of either solution can have a huge effect on the pH at the end point. i.e. one drop of acid can drop the pH by several units.

- any indicator that changes colour in the range of 3 – 11 can be used in a strong acid – strong base titration.

- indicators like phenolphthalein are very good because they have very distinct differences between the colours in acid and base and so the end-point can be easily seen.

Exercise 3:

Below are some questions to assist your understanding. Complete Exercise 3 and check your answers at the end of the book.
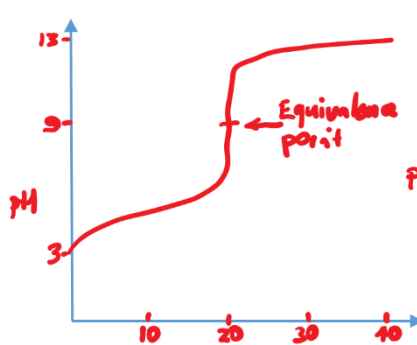
Select an indicator for the following titrations and explain your reasons.

a) Na$_2$CO$_3$ and HCl

b) H$_2$CO$_3$ and NaOH

c) methyl orange is not used in weak acid – strong base titration

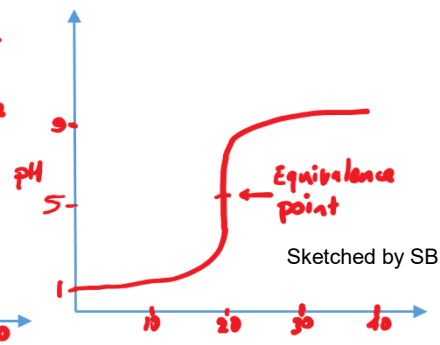d) phenolphthalein is used for weak acid – strong base titrations and for strong acid – strong base titrations.
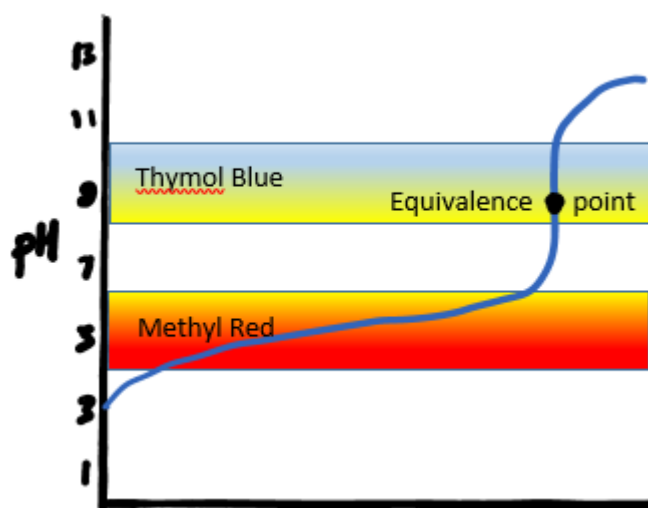
Titration Curves
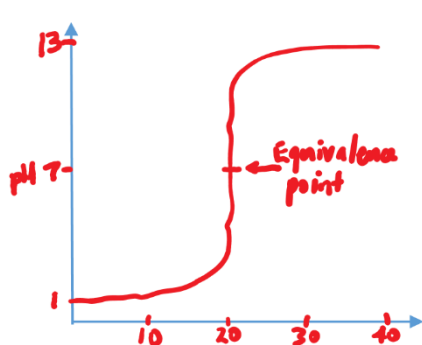
| Strong acid – strong base. 20.0 mL of 0.10 mol L^{-1} HCl with 0.10 mol L^{-1} NaOH | Weak acid – strong base. 20.0 mL of 0.10 mol L^{-1} CH$_3$COOH with mol L^{-1} NaOH | Strong acid – weak base. 20.0 mL of 0.10 mol L^{-1} HCl with 0.10 mol L^{-1} NH$_3$ |

Sketched by SB

Choosing the correct indicator is very important. For this weak acid – strong base titration, thymol blue would be a good choice as the colour change (end point) occurs at the equivalence point.

If methyl red was selected for this titration, the end point would occur before the reaction was complete.

A lower volume of base will be recorded and as a result, calculations will be incorrect.

Exercise 4: Titrations

Below are some questions to assist your understanding. Complete Exercise 4 and check your answers at the end of the book.

1.   For each of the following titrations identify:

   (i)   The substance present at the equivalence point (write the equation first).

   (ii)   Whether the solution at the equivalence point is acidic, basic or neutral.

   (iii) An appropriate indicator for the titration.

a.   The titration of hydrofluoric acid with sodium hydroxide solution.

b.   The titration of potassium hydroxide with nitric acid.

c.   The titration of sodium carbonate with hydrochloric acid.

2.   In the titration of ethanoic acid with sodium hydroxide solution, phenolphthalein is an appropriate indicator. If methyl orange was used as the indicator, what effect would this have on the apparent volume of sodium hydroxide required for neutralisation?

3.   Explain why the following substances are not suitable for use as a primary standard?

a.   sodium hydroxide

b.   hydrochloric acid

4.   A student carrying out a titration between HCl acid and ammonia used phenolphthalein as an indicator.

a.   Is this an appropriate indicator?

b.   Explain clearly why/why not.

5.   A student is performing a titration and has a 25.0 mL solution that is at pH of 9.0 due to a higher $[OH^-]$ than $[H^+]$.

Calculate the $[H^+]$.

Calculate the number of moles of OH^-.

1 drop of 1 M HCl is added. Assuming the volume of 1 drop is 0.05 mL, calculate the new pH.
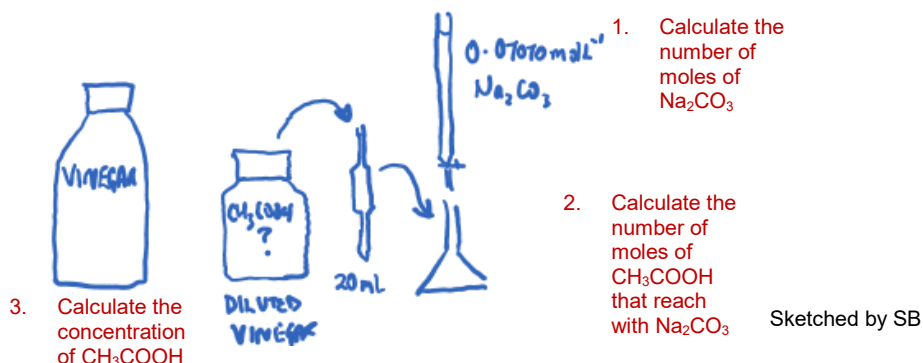
Data obtained from acid-base titrations can be used to calculate the masses of substances and concentrations and volumes of solutions involved.

Acid-Base Calculations in volumetric analysis

Find the Concentration of a Solution from Titration Data

Example: Determine the concentration of ethanoic acid in a diluted vinegar solution if 20.00 mL portions of diluted vinegar are titrated with 7.070×10^{-2} mol L^{-1} sodium carbonate solution. Several titrations were performed and the volumes of sodium carbonate solution were 23.55 mL, 22.85 mL, 22.70 mL and 22.80 mL

It is often useful to sketch a diagram of the situation.

1. Calculate the number of moles of Na_2CO_3

2. Calculate the number of moles of CH_3COOH that reach with Na_2CO_3 Sketched by SB

3. Calculate the concentration of CH_3COOH

Find av. titration vol. Discard 23.55	V(average) = $\dfrac{22.85 + 22.70 + 22.80}{3}$ = 22.78 mL
Find the av. moles of Na_2CO_3	$n(Na_2CO_3) = cV = 7.070 \times 10^{-2} \times 22.78 \times 10^{-3}$ = 1.611×10^{-3} mol
2 moles react with 1 mole	$2CH_3COOH + Na_2CO_3 \rightarrow 2\,NaCH_3COO + H_2O + CO_2$
Calculate moles of CH_3COOH	$n(CH_3COOH) = 2 \times n(Na_2CO_3) = 2 \times 1.611 \times 10^{-3}$ = 3.222×10^{-3} mol
Calculate conc. of CH_3COOH	$c(CH_3COOH) = \dfrac{n}{V} = \dfrac{3.22 \times 10^{-3}}{20 \times 10^{-3}}$ = 1.61×10^{-1} mol L^{-1}

Exercise 5:

Below are some questions to assist your understanding. Complete Exercise 5 and check your answers at the end of the book.

1. a) Calculate the concentration of a solution of sodium carbonate made when 16.56 g is dissolved in distilled water and the solution is made up to 2.000 L.

    b) This solution was used to standardise a solution of hydrochloric acid, which was known to have a concentration of approximately 0.1 mol L^{-1}. It was found that 12.95 mL of the sodium carbonate solution was required to neutralise 20.00 mL f the hydrochloric acid. Calculate the exact concentration of the hydrochloric acid.

c)   This hydrochloric acid was then used to calculate the concentration of sodium hydroxide in a household cleaning solution. 10.00 mL samples of the cleaning solution were diluted by adding distilled water up to a volume of 250.0 mL in a volumetric flask. The average value of the hydrochloric acid required to neutralise 20.00 mL aliquots of the diluted cleaner was 17.75 mL. Calculate the concentration of the sodium hydroxide in the original cleaning solution.

2. In a titration 0.105 mol L^{-1} HCl is used to standardise KOH solution using phenolphthalein as an indicator. 21.1 mL of the acid is needed to neutralise 25.0 mL of KOH solution. What is the concentration of the KOH solution?

3. The concentration of an unknown sodium carbonate solution was to be determined by titration with 1.00 M nitric acid using methyl orange indicator. 3.5 mL of nitric acid was added to 25.0 mL of sodium carbonate when a colour change from yellow to red indicated the end-point of the titration. What is the concentration of the sodium carbonate solution?

4. 24.4 mL of HCl is needed to neutralise 25.0 mL of 0.104 M NaOH solution. Calculate:

  a) the concentration of HCl in mol L^{-1} and g L^{-1}
  b) the volume of the original acid solution needed to make 1.00 mL of 0.100 M solution and
  c) the volume of water which must be added to 2.00 L of the original acid solution to make it 0.100 mol L^{-1}.

5. A sample of vinegar has a density of 1.01 g mL^{-1} and contains 3.00 % by mass of ethanoic (acetic) acid. What volume of 0.500 M potassium hydroxide is required to neutralise 25.0 mL of the vinegar?

Titration Calculation Involving a Dilution

Example:

Prior to analysis, a vinegar solution is diluted by pipetting 50.00 mL of it into a 250 mL volumetric flask and making this up to the mark with distilled water.
A 20.00 mL sample of diluted vinegar required on average 23.65 mL of 7.070 x 10^{-2} mol L^{-1} sodium carbonate solution to reach the end point. Determine the concentration of ethanoic acid in the undiluted vinegar solution.

Sketched by SB
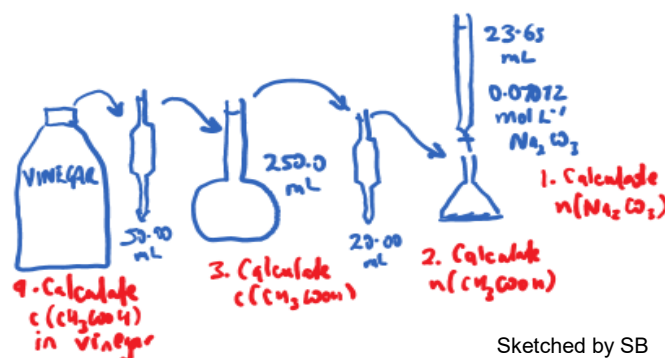
|	
Find the av. moles of Na_2CO_3	$n(Na_2CO_3) = cV = 7.070 \times 10^{-2} \times 23.65 \times 10^{-3} = 1.672 \times 10^{-3}$ mol
2 moles react with 1 mole	$2CH_3COOH + Na_2CO_3 \rightarrow 2\ NaCH_3COO + H_2O + CO_2$
Calculate moles of CH_3COOH	$n(CH_3COOH) = 2 \times n(Na_2CO_3) = 2 \times 1.672 \times 10^{-3}$ $= 3.344 \times 10^{-3}$ mol
Calculate conc. of CH_3COOH	$c(CH_3COOH) = \dfrac{n}{V}_{diluted} = \dfrac{3.\,344 \times 10^{-3}}{20 \times 10^{-3}} = 1.672 \times 10^{-1}$ mol L^{-1}
Find original conc. of CH_3COOH	$c_o \times v_o = c_n \times v_n \qquad\qquad c_o = \dfrac{c_n \times v_n}{v_o}$

$$c_o = \frac{1.672 \times 10^{-1} \times 250 \times 10^{-3}}{50 \times 10^{-3}} = 0.836 \text{ mol } L^{-1}$$

- **Another method without calculating the concentration of CH_3COOH in diluted solution:**

$$2CH_3COOH + Na_2CO_3 \rightarrow 2NaCH_3COO + H_2O + CO_2$$
$$2 \text{ mol} \qquad\qquad 1 \text{ mol}$$
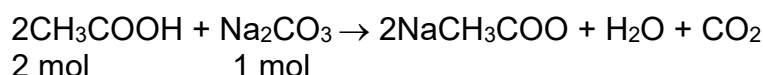
thus $n(CH_3COOH$ in 20.00 mL aliquot$)\ =\ 2 \times n(Na_2CO_3) = 3.344 \times 10^{-3}$ mol.

$n(CH_3COOH$ in 250 mL of diluted solution$) = \dfrac{250}{20.00} \times 3.344 \times 10^{-3}$

$$= 0.0418 \text{ mol}$$

$n(CH_3COOH$ in original 50.00 mL$) = n(CH_3COOH$ in 250 mL of dilute solution

$$= 0.0418 \text{ mol}$$

$c(CH_3COOH$ in original 50.00 mL$) = \dfrac{n}{V} = \dfrac{0.0418}{0.050}$

$$= 0.836 \text{ mol } L^{-1}$$

Exercise 6:

Below are some questions to assist your understanding. Complete Exercise 6 and check your answers at the end of the book.

1. An analyst has a pure substance 'A' which is a weak acid and she decides to determine its molecular weight by titration.
   She dissolves 1.3867 g of 'A' in water and makes the solution up to 250.0 mL in a volumetric flask. She then titrates 20.00 mL portions of the solution 'A' with 0.09826 mol L^{-1} sodium hydroxide *from the burette* using phenolphthalein as indicator, and records the following titration figures.

Initial reading (mL)	20	19.5	18.32	17.77	18.87
Final reading (mL)	0.11	1.32	0.67	0.08	1.20

   (a) Calculate the appropriate value for the volume of sodium hydroxide required for titration.

   (b) Assume that 'A' is monoprotic acid and calculate the number of moles of 'A' in the 1.3867 g of 'A'.

   (c) Use this value to calculate the molecular weight of 'A'.

   (d) Assume that 'A' is a diprotic acid and calculate its molecular weight.

   (e) Assume that 'A' is a triprotic acid and calculate its molecular weight.

2. A 10.00 mL sample of sulfuric acid (H_2SO_4) is taken from a car battery. This sample is made up to 250.0 mL with distilled water. The diluted acid is placed in a burette and reacted with 20.00 mL aliquots of sodium hydroxide solution, which has a concentration of 0.203 mol L^{-1}. Phenolphthalein was used as the indicator and on average 14.25 mL of the sodium hydroxide was required for neutralisation.

Back Titrations

A back titration, or indirect titration, is generally a two-stage analytical technique.

Reactant A of unknown concentration is reacted with excess reactant B of known concentration.

A titration is then performed to determine the amount of reactant B in excess. From calculations, the amount of reactant B used can be determined and then the concentration of reactant A can be determined. Click on the following link to learn more about titrations. https://www.ausetute.com.au/titrcalc.html.

Back titrations are used in a number of situations. For example;

- o one of the reactants is volatile, for example ammonia

- o an acid or a base is an insoluble solid, for example calcium carbonate
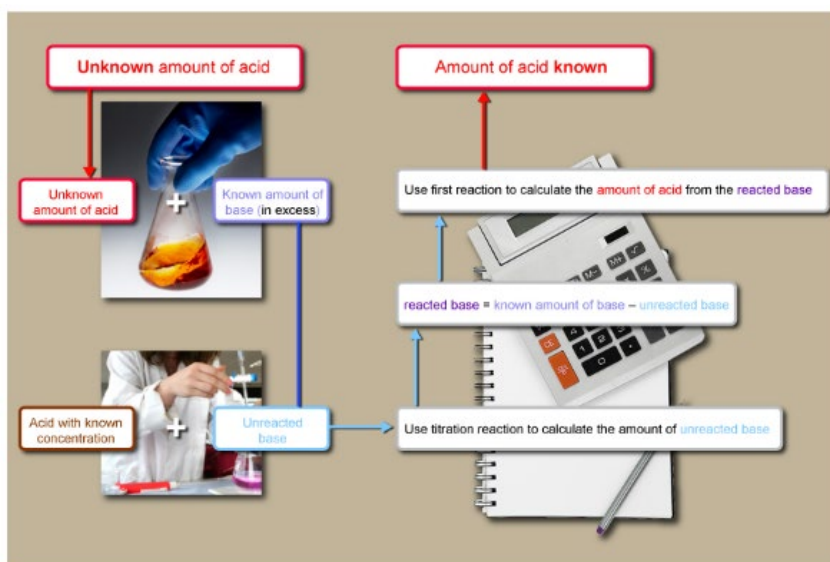
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- o a particular reaction is too slow

- o direct titration would involve a weak acid - weak base titration

- o the end-point of this type of direct titration is very difficult to observe.

Example: A 2.50 g sample of limestone was dissolved in 75.0 mL of 0.996 mol L^{-1} hydrochloric acid. After heating to remove all the CO_2, the solution was cooled and diluted to 250 mL. A 25.0 mL sample of the diluted solution was titrated with 25.27 mL of a standardised of 0.1215 mol L^{-1} sodium hydroxide. Determine the percentage of calcium carbonate in the original limestone sample.

$NaOH + HCl \rightarrow NaCl + H_2O$
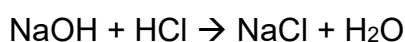
n(NaOH) = cV = 0.1215 x 0.02527 = 0.003070 mol

n(HCl 25 mL sample) = 0.003070 mol

n(HCl 250 mL) = 10 x 0.003070 = 0.03070 mol

n(HCl original) = cV = 0.996 x 0.0750 = 0.0747 mol

n(HCl reacted) = 0.0747 – 0.03070 = 0.0440 mol

$2HCl + CaCO_3 \rightarrow CaCl_2 + CO_2 + H_2O$

n(CaCO$_3$) = ½ n(HCl) = 0.0220 mol

m(CaCO$_3$) = n x M
= 0.0220 x 100.09 = 2.201 g

%(CaCO$_3$) = 2.201/2.50 x 100 = 88.1%

Exercise 7:

Below are some questions to assist your understanding. Complete Exercise 7 and check your answers at the end of the book.

1. A chemist wanted to calculate the mass of citric acid ($C_6H_8O_7$) in a lemon. (Citric acid is a triprotic acid.) All the juice from one lemon was mixed with 40.00 mL of a 0.500 mol L^{-1} sodium hydroxide solution.

   The resulting solution was filtered and titrated against hydrochloric acid with a concentration of 1.05 mol L^{-1} using methyl orange as an indicator. It was found that 15.90 mL of the acid was required for the titration. Click on the steps below to solve the problem.

2. An experiment was carried out to find the purity of chalk, which comprises of calcium carbonate ($CaCO_3$). A 3.63 g sample of impure chalk was added to 50.0 mL of 2.00 mol L^{-1} HCl. The resulting solution was filtered into a volumetric flask and made up to 250.0 mL.
   25.00 mL aliquots of this solution were then titrated against 0.105 mol L^{-1} NaOH solution and the results shown below:

Burette readings (mL)	Titrations		
	1	2	3
Final volume	32.50	37.25	43.15
Initial volume	0.00	5.50	11.30
Titre			

a) Complete the table and calculate the titration volume.

b) Calculate the number of moles of hydrochloric acid present in the 25.00 mL aliquots.

c) Calculate the total number of moles of hydrochloric acid present in the 250.0 mL flask and hence calculate the % purity of the chalk.

Acid and Base Titrations Questions

1.   What is a titration?

2.   What is the essential difference between a pipette and a burette?

3.   What are the features of a primary standard? Name one primary standard used in acid/base titrations.

4.   What is wrong with using NaOH as a primary standard?

5.   Why are indicators used in acid/base titrations?

6.   Distinguish between equivalence and end point.

7.   Draw pH versus volume added graphs for:

   (i)   sodium carbonate solution + hydrochloric acid (acid in burette)

   (ii)   HCl solution vs NaOH solution (acid in burette)

   (iii) CH_3COOH solution vs NaOH solution (acid in burette)

   On the graphs mark the equivalence point.

8.   Write equations explaining why the pH at equivalence is what is shown on the graphs in question 7.

9.   Suggest appropriate indicators for the titrations above. On the graphs, mark the range in which their colour changes.

10.   Predict why weak acid/weak base titrations are experimentally difficult to perform.

Acid – Base Neutralisation Questions

1. Calculate the number of moles of sodium hydroxide that would be required to react with ALL the hydrogen ions from:

    i.   2.00 moles of hydrochloric acid.
    ii.   1.00 mole of sulfuric acid.
    iii.   0.500 moles of phosphoric acid.
    iv.   1.40 moles of nitric acid.
    v.   0.200 moles of sulfuric acid.

2. Determine how many mL of 3.00 mol L^{-1} sodium hydroxide are required to neutralise 30.0 mL of 2.00 mol L^{-1} hydrochloric acid solution.

3. 2.00 g of sodium hydroxide was dissolved in water and the solution made up to 2.50×10^2 mL Determine:

    i.   the molarity of the solution
    ii.   the strength of the solution in g L^{-1}
    iii.   the volume of 0.100 mol L^{-1} sulfuric acid required to neutralise 25.0 mL of the solution
            and
    iv.   the mass of pure sulfuric acid required for complete neutralisation.

4. Calculate the volume of 0.400 mol L^{-1} sodium hydroxide which is required to completely neutralise:

    i.   20.0 mL of 1.00 mol L^{-1} nitric acid.
    ii.   15.0 mL of 0.200 mol L^{-1} sulfuric acid.
    iii.   50.0 mL of 6.00 x 10^{-2} mol L^{-1} oxalic acid.

5. A solution of hydrochloric acid (HCℓ) is prepared by diluting 50.0 mL of 5.00 mol L^{-1} acid to 2.50 x 10^2 mL in a volumetric flask. What volume of 0.400 mol L^{-1} sodium hydroxide is required to neutralise 40.0 mL of the acid solution?

6. A solution of crystalline oxalic acid ($H_2C_2O_4.2H_2O$) is prepared by dissolving 12.6 g of the acid in 4.00 L of water. 20.0 mL of this solution are required to completely neutralise 36.0 mL of a potassium hydroxide solution. Calculate the molarity and strength in g L^{-1} of the potassium hydroxide solution.

7. A solution of sulfuric acid is prepared from a 98% by mass concentrated sulfuric acid solution by diluting 20.0 g of the acid to 2.50×10^2 mL. Determine the volume of 0.110 mol L^{-1} of sodium carbonate required to completely neutralise 10.0 mL of the diluted acid solution.

8. A 100.0 mL sample of a 1.00 mol L^{-1} sodium carbonate solution is diluted to 2.50×10^2 mL. 20.0 mL of this diluted solution was reacted with 4.00 g of a 6.50% by mass solution of hydrochloric acid. 1.00 g of the hydrochloric acid solution has a volume of 0.976 mL. Determine the volume of acid required to complete the reaction.

9. Determine the volume of 0.125 mol L^{-1} ammonia solution required to neutralise 60.0 mL of 0.200 mol L^{-1} nitric acid solution.

10. Calculate the molarity of a solution of sulfuric acid if 25.0 mL of the acid dissolves 1.00 g of magnesium metal.

11. 10.0 mL of concentrated nitric acid was diluted to 2.50 × 10^2 mL.   25.0 mL of this diluted solution required 40.0 mL of 5.00 × 10^{-2} mol L^{-1} sodium hydroxide solution for complete neutralisation. Determine the molarity of the original concentrated form of nitric acid.

12. 3.00 g of a monoprotic acid was dissolved and made up to 2.50 × 10^2 mL.   20.0 mL of this prepared solution was found to require 30.0 mL of 0.100 mol L^{-1} sodium hydroxide solution for complete neutralisation. Determine the formula mass of the acid.

13. 2.00 g of a triprotic acid was dissolved in distilled water and then made up to 2.50 × 10^2 mL. 40.0 mL of this acid solution required 50.0 mL of 0.100 mol L^{-1} potassium hydroxide solution for complete neutralisation. Determine the formula mass of the acid.

14. 1.00 x 10^2 mL of 0.100 mol L^{-1} sulfuric acid is treated with 2.00 × 10^2 mL of 0.500 mol L^{-1} sodium hydroxide solution.   Complete neutralisation is completed by the addition of   150.0 mL of sulfuric acid. Calculate the molarity of the acid.

15. 50.0 ml of 0.200 mol L^{-1} sulfuric acid is treated with 100.0 mL of 0.250 mol L^{-1} sodium hydroxide solution. Complete neutralisation was completed with addition of 75.0 mL of sulfuric acid. Calculate the molarity of the sulfuric acid which was added to complete the neutralisation.

16. Washing soda has the chemical formula of $Na_2CO_3.XH_2O$.   2.75 g of washing soda was dissolved in distilled water and made up to 300.0 mL. 50.0 mL of this solution required
25.0 mL of 6.30 × 10^{-2}mol L^{-1} sulfuric acid for complete neutralisation. Determine the value of 'X'.

17. A sample of sodium hydrogencarbonate was heated until a constant mass was recorded. This sample was then dissolved in water and the solution made up to 1.00 L. 25.0 mL of this solution required 11.25 mL of a 0.100 mol L^{-1} sulfuric acid for complete neutralisation. Calculate the mass of the original sample of sodium hydrogencarbonate.

18. 0.700 g of an ammonium salt was dissolved in distilled water and then 25.0 mL of a 1.10 mol L^{-1} sodium hydroxide solution was added. The solution was boiled. All ammonia gas was expelled. After cooling it was found that 30.0 mL of 0.450 mol L^{-1} hydrochloric acid was required to neutralise the excess base. Determine the percentage of ammonia in the salt.

19. To neutralise 10.0 mL of hydrochloric acid, 20.0 mL of 2.00 mol L^{-1} sodium hydroxide solution was required. In another experiment, 10.0 mL of the same acid dissolved

1.00 g of calcite. 10.25 mL of 2.00 mol L^{-1} of sodium hydroxide solution was required to neutralise the excess hydrochloric acid. Calculate the percentage of calcium carbonate in the calcite sample.

20.   Determine the percentage purity of a sample of zinc if 0.325 g required 38.0 mL of 0.250 mol L^{-1} hydrochloric acid for a complete reaction.

21.   An approximately 0.100 mol L^{-1} solution of sulfuric acid was standardised by titration against massed quantities of anhydrous sodium carbonate. It was found that 0.200 g of sodium carbonate required 39.2 mL of the acid to complete neutralisation. Determine the molarity of the acid.

22.   0.0920 g of sodium metal is added to water. When the reaction has completed the whole solution is treated with a 0.250 mol L^{-1} hydrochloric acid solution. Determine the volume of acid required for complete neutralisation. Name an indicator which could be used.

23.   0.400 g of a divalent metal is dissolved in 100.0 mL of 0.500 mol L^{-1} hydrochloric acid.
68.0 mL of 0.250 mol L^{-1} sodium hydroxide solution was required to neutralise the excess acid. Determine the relative atomic mass of the metal.

24.   0.100 g of calcium carbonate was dissolved in 25.0 mL of hydrochloric acid solution, and it was found that 5.50 mL of sodium hydroxide solution were required to neutralise the excess acid. A second titration showed that 27.5 mL of sodium hydroxide neutralised 25.0 mL of the hydrochloric acid solution. Determine the molarity of the sodium hydroxide and hydrochloric acid solutions.

Answers

Exercise 1: Standard Solutions

a) $n(Na_2CO_3) = \dfrac{m}{M} = \dfrac{1.336}{105.99} = 0.01260$ mol

$\qquad$ 4 sf

$\quad c(Na_2CO_3) = \dfrac{n}{V} = \dfrac{0.0126}{0.250} = 5.041 \times 10^{-2}$ mol L^{-1}

b) $n(H_2C_2O_4.2H_2O) = cV = 0.250 \times 0.500 = 0.125$ mol

$\quad m(H_2C_2O_4.2H_2O) = nM = 0.125 \times 126.068 = 15.8$ g $\qquad M(H_2C_2O_4.2H_2O) =$
126.068

c) $CH_3COOH + NaOH \rightarrow CH_3COONa + H_2O$

$\quad n(NaOH) = cV = 0.150 \times 0.02450 = 0.003675$ mol

$\quad n(CH_3COOH) = \dfrac{1}{1} \times n(NaOH) = 0.003675$ mol

$\quad c = \dfrac{n}{V} = \dfrac{0.003675}{0.005} = 0.735$ mol L^{-1}

Exercise 2: Acid Base Reactions.
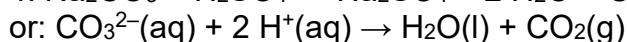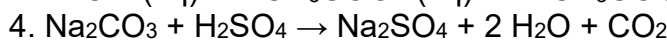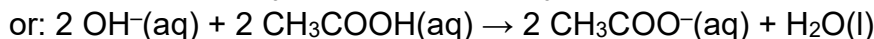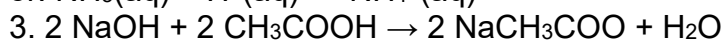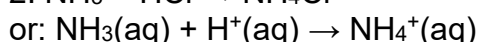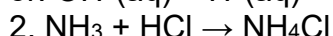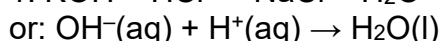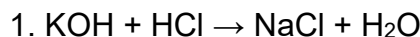
1. $KOH + HCl \rightarrow NaCl + H_2O$
or: $OH^-(aq) + H^+(aq) \rightarrow H_2O(l)$
2. $NH_3 + HCl \rightarrow NH_4Cl$
or: $NH_3(aq) + H^+(aq) \rightarrow NH_4^+(aq)$
3. $2\ NaOH + 2\ CH_3COOH \rightarrow 2\ NaCH_3COO + H_2O$
or: $2\ OH^-(aq) + 2\ CH_3COOH(aq) \rightarrow 2\ CH_3COO^-(aq) + H_2O(l)$
4. $Na_2CO_3 + H_2SO_4 \rightarrow Na_2SO_4 + 2\ H_2O + CO_2$
or: $CO_3^{2-}(aq) + 2\ H^+(aq) \rightarrow H_2O(l) + CO_2(g)$
pH
1. The products are Na^+, Cl^- and H_2O.
$\quad$ Na^+ and Cl^- will not hydrolyse water so the solution is neutral – pH ~7.
2. Product is NH_4^+.
$\quad$ $NH_4^+ + H_2O \rightarrow NH_3 + H_3O^+$
$\quad$ This will create an acidic solution – pH below 7.
3. Product are CH_3COO^- and H_2O
$\quad$ $CH_3COO^- + H_2O \rightarrow CH_3COOH + OH^-$
$\quad$ This will create a basic solution – pH above 7.
4. Products are CO_2 and H_2O.
$\quad$ These combine to form H_2CO_3 which releases H^+ into solution.
$\quad$ This will create an acidic solution – pH below 7.
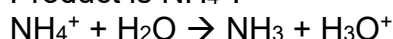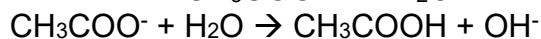
Exercise 3:

a) methyl orange (SA + WB)
b) Phenolphthalein (WA + SB)
c) Methyl orange changes colour between a pH
of 3.1–4.4 which is on the acid side. A weak
acid and strong base forms a basic solution.
The end point of methyl orange is not inside
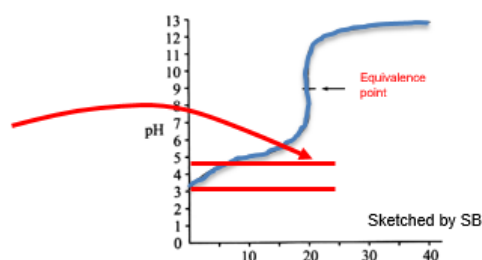
this reaction's equivalence point. Colour change (end point) will occur before the equivalence point.

d) A weak acid-strong base reaction forms a basic solution and a strong acid-strong base forms a neutral solution. Phenolphthalein end point will match these equivalence points.
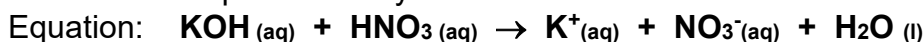
Exercise 4: Titrations
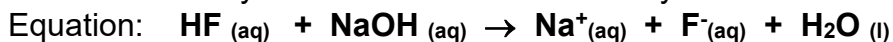
1.
   a. The titration of hydrofluoric acid with sodium hydroxide solution.
      Equation: $HF_{(aq)} + NaOH_{(aq)} \rightarrow Na^+_{(aq)} + F^-_{(aq)} + H_2O_{(l)}$
      Substance present: $Na^+_{(aq)}$; $F^-_{(aq)}$; $H_2O_{(l)}$
      Acidic or basic: **basic due to $F^-_{(aq)}$**   Indicator: **phenolphthalein**
   b. The titration of potassium hydroxide with nitric acid.
      Equation: $KOH_{(aq)} + HNO_{3\,(aq)} \rightarrow K^+_{(aq)} + NO_3^-_{(aq)} + H_2O_{(l)}$
      Substance present: K^+ ; NO_3^- ; H_2O
      Acidic or basic: **Neutral**   Indicator: **phenolphthalein, methyl red, litmus**
   c. The titration of sodium carbonate with hydrochloric acid.
      Equation: $Na_2CO_{3\,(s)} + HCl_{(aq)} \rightarrow Na^+_{(aq)} + Cl^-_{(aq)} + CO_{2\,(g)} + H_2O_{(l)}$
      Substance present: $Na^+_{(aq)}$; $Cl^-_{(aq)}$; $CO_{2\,(g)}$; $H_2O_{(l)}$
      Acidic or basic: **acidic due to CO_2**   Indicator: **methyl orange**
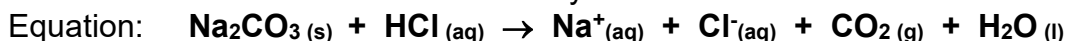
2.   **The solution at equivalent point would contain sodium ethanoate – an alkaline solution. Methyl orange is suitable for acidic solutions, so if methyl orange was used it would change colour (end point) before the equivalence point was reached. Therefore the volume added would be less than required to reach equivalence point.**

3.
   a. **difficult to obtain pure**
      deliquescent (absorbs water)
   b. **not pure, results from dissolving an approximate amount of gas in solution.**

4.
   a. **No**
   b. **Neutralisation produces an acidic end point due to presence of ammonium ion. Phenolphthalein only suitable for basic and neutral equivalence points.**

5.
   $[H^+] = 10^{-pH} = 10^{-9} = 1 \times 10^{-9}$ mol L^{-1}
   Calculate the number of moles of OH^-.
   $[OH^-] = (1 \times 10^{-14})/(1 \times 10^{-9}) = 1 \times 10^{-5}$ mol L^{-1} **pOH = 5   [OH⁻] = 1 x 10⁻⁵ M**
   $n(OH^-) = cV = 1 \times 10^{-5} \times 0.025 = 2.5 \times 10^{-7}$ mol $n(OH^-) = 2.5 \times 10^{-7}$ mol
   1 drop of 1 M HCl is added. Assuming the volume of 1 drop is 0.05 mL, calculate the new pH.
   $n(H^+) = cV = 1 \times 0.00005 = 0.00005 = 5.0 \times 10^{-5}$ mol
   excess $H^+ = n(H^+) - n(OH^-) = 5.0 \times 10^{-5} – 2.5 \times 10^{-7} = 4.975 \times 10^{-5}$ mol   c=n/V = 1.99x10⁻³
   pH = $-\log_{10}[H^+]$ = $-\log_{10}(1.99 \times 10^{-3})$ = 2.7

Exercise 5:

1.

(a) $n(Na_2CO_3) = = \frac{m}{M} = \frac{16.56}{105.99} = 0.1562$ mol

$c(Na_2CO_3) = \frac{n}{V} = \frac{0.1562}{2.000} = 0.0781$ mol L^{-1}
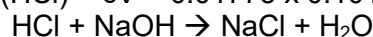
(b) $Na_2CO_3 + 2\ HCl \rightarrow 2\ NaCl + H_2O + CO_2$
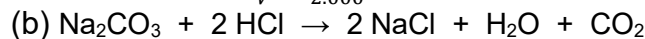
$n(Na_2CO_3) = c \times V = 0.0781 \times 0.012195 = 1.011 \times 10^{-3}$ mol

$n(HCl) = \left(\frac{2}{1}\right) \times n(Na_2CO_3) = \left(\frac{2}{1}\right) \times 1.011 \times 10^{-3} = 2.023 \times 10^{-3}$

$c(HCl) = \left(\frac{n}{V}\right) = \frac{2.023\ x\ 10^{-3}}{0.02000} = 0.101$ mol L^{-1}

(c) $n(HCl) = cV = 0.01775\ x\ 0.1011661 = 0.0017956$ mol

$HCl + NaOH \rightarrow NaCl + H_2O$

$n(NaOH) = n(HCl) = 0.0017956$ mol

$c(NaOH) = n/V = 0.0017956/0.0200 = 0.0897849$ mol L^{-1} (in 20 mL pipette and 250 mL flask)

$c(NaOH\ in\ 10mL\ pipette) = 0.0897849\ x\ 250/10 = 2.2446$ mol L^{-1}

which is the concentration in the original sample = 2.245 mol L^{-1}

2. $H^+ + OH^- \rightarrow H_2O$

$n(HCl) = c\ V = $ **0.105 x 21.1 x 10^{-3}**

$= $ **2.2155 x 10^{-3} mol**

1 mol H$^+$ reacts with 1 mol OH$^-$

$n(KOH) = $ **2.2155 x 10^{-3} mol**

$c(KOH) = n/V = $ **$\frac{2.2155\ x\ 10^{-3}}{25.0\ x\ 10^{-3}}$** $= $ **8.86 x 10^{-2} mol L^{-1}**

3. $2H^+ + CO_3^{2-} \rightarrow CO_2 + H_2O$

$n(HNO_3) = c\ V = $ **1.00 x 3.5 x 10^{-3}**
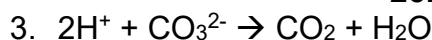
$= $ **3.5 x 10^{-3} mol**

2 mol H$^+$ reacts with 1 mol CO$_3^{2-}$

$n(Na_2CO_3) = $ **½ x 3.5 x 10^{-3}** $= $ **1.75 x 10^{-3} mol**

$c(Na_2CO_3) = n/V$

$= $ **$\frac{1.75\ x\ 10^{-3}}{25.0\ x\ 10^{-3}}$**

$= $ **7.00 x 10^{-2} mol L^{-1}**

4.

(a) $n(NaOH) = c\ V = $ **0.104 x 25.0 x 10^{-3}**

$= $ **2.60 x 10^{-3} mol**

1 mol H$^+$ reacts with 1 mol OH$^-$

$n(HCl) = $ **2.60 x 10^{-3} mol**

$c(HCl) = n/V$

$= $ **$\frac{2.60\ x\ 10^{-3}\ mol}{24.4\ x\ 10^{-3}}$**

$= $ **1.07 x 10^{-1} M**

$n = m/M$

$m = n\ M$

$= $ **1.07 x 10^{-1} x 36.458**

$= $ **3.88** g in **1** L

$= $ **3.88** gL^{-1}

g L^{-1} = mol L^{-1} x M

mol L^{-1} = $\frac{M}{g\ L^{-1}}$

(b) $c_1V_1 = c_2V_2$

$V_1 = $ **$\frac{c_2V_2}{c_1}$** $= $ **$\frac{0.100\ x\ 1\ x\ 10^{-3}}{1.07\ x\ 10^{-1}}$**

$= $ **9.38 x 10^{-4} L**

(c) $c_1V_1 = c_2V_2$

$$V_2 = \frac{c_1 V_1}{C_2} = \frac{1.07 \times 10^{-1} \times 2.00}{0.100}$$

$$= 2.14 \text{ L} \quad \text{i.e. add } 140 \text{ mL}$$

5. 3.00 % of 1.01 g = **3.03 x 10^{-2}** g mL^{-1} ethanoic acid CH_3COOH

$$n\,(CH_3COOH) = m/M$$

$$= \frac{3.03 \times 10^{-2} \times 25}{60.032}$$

$$= \frac{7.575 \times 10^{-1}}{60.032} \qquad = 1.2618 \times 10^{-2} \text{ mol}$$

1 mol H$^+$ reacts with 1 mol OH$^-$

n(KOH) = **1.2618 x 10^{-2} mol**

V(KOH) = n/c

$$= \frac{1.2618 \times 10^{-2}}{0.500}$$

$$= 2.524 \times 10^{-2} \text{ L}$$

$$= 25.2 \text{ mL}$$

Exercise 6:

1.

Initial reading (mL)	20	19.5	18.32	17.77	18.87
Final reading (mL)	0.11	1.32	0.67	0.08	1.20
	19.89	**18.18**	**17.65**	**17.69**	**17.67**

a) $\dfrac{(17.65 + 17.69 + 17.67)}{3}$ = **17.67 mL**   **(Discard the first two values)**

b) n(NaOH) = cV = 0.09826 x 17.67 x 10^{-3} = **1.74 x 10^{-3} mol**
   n(A) = n(NaOH) = **1.74 x 10^{-3} mol** in 20 mL = **0.0217 mol** in 250 mL

c) M(A) = $\dfrac{m}{n}$ = $\dfrac{1.3867}{0.0217}$ = **63.8 g mol^{-1}**

d) n(A) = ½ x 0.0217 = **0.0109 mol**
   M(A) = $\dfrac{m}{n}$ = $\dfrac{1.3867}{0.0109}$ = **127.6 g mol^{-1}**

e) n(A) = ⅓ x 0.0217 = **0.00725 mol**
   M(A) = $\dfrac{m}{n}$ = $\dfrac{1.3867}{0.00725}$ = **191.4 g mol^{-1}**

2. Step 1: Calculate the number of moles of sodium hydroxide

$$n(NaOH) = c \times V = 0.203 \times 0.01425 = 2.89 \times 10^{-3} \text{ mol}$$

Step 2: Use the equation to work out the moles of acid present

$$2\,NaOH + H_2SO_4 \rightarrow Na_2SO_4 + H_2O$$
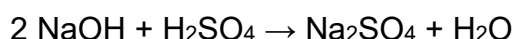
$$n(H_2SO_4) = \left(\frac{1}{2}\right) \times n(NaOH)$$

$$n(H_2SO_4) = \left(\frac{1}{2}\right) \times 2.89 \times 10^{-3} = 1.45 \times 10^{-3} \text{ mol}$$

Step 3: Calculate the concentration of the diluted acid

$$c(H_2SO_4) = \frac{n}{V} = \frac{1.45 \times 10^{-3}}{0.0200} = 7.25 \times 10^{-2} \text{ mol L}^{-1}$$

Step 4: Calculate the concentration of the original acid

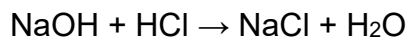$$c(H_2SO_4) = \left(\frac{250.0}{10.00}\right) \times 7.25 \times 10^{-2} = 1.81 \text{ mol L}^{-1}$$

Exercise 7

1.

Step 1: Calculate the moles of acid (HCl) in the titration

$n(HCl)_{titration} = c \times V = 1.05 \times 0.01590 = 0.0167$

Step 2: Equation for titration reaction

$NaOH + HCl \rightarrow NaCl + H_2O$

Step 3: Calculate the moles of base (NaOH) used in the titration

$n(NaOH)_{titration} = n(HCl)_{titration} = 0.0167 \text{ mol}$

(This is the number of moles of base that was *not* used up in the first reaction.)

Step 4: Calculate the total moles of base in the original solution

$n(NaOH)_{original} = c \times V = 0.500 \times 0.0400 = 0.0200 \text{ mol}$

Step 5: Calculate the total moles of base used in the first reaction

$n(NaOH)_{first\ reaction} = 0.0200 - 0.0167 = 3.30 \times 10^{-3} \text{ mol}$

Step 6: Write the equation for the first reaction

$C_6H_8O_7 + 3\ NaOH \rightarrow Na_3C_6H_5O_7 + 3\ H_2O$
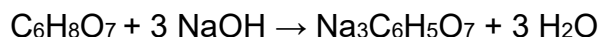
Step 7: Calculate the moles of acid used in the first reaction

(Note: the acid used in the first reaction was $C_6H_8O_7$)

$$n(C_6H_8O_7) = \left(\frac{1}{3}\right) \times n(NaOH) = \left(\frac{1}{3}\right) \times 0.0033 = 0.00110$$

Step 8: Calculate the mass of acid in the original sample

(Note: the acid used in the first reaction was $C_6H_8O_7$)

$m(C_6H_8O_7) = n \times M = 0.00110 \times 192.124 = 0.211$ g

2.

Titre	32.50	31.75	31.85

a)   Complete the table and calculate the titration volume.

$$V(NaOH) = \frac{(31.75 + 31.85)}{2} = 31.80 \text{ mL} = 0.03180 \text{ L}$$

b)   Calculate the number of moles of hydrochloric acid present in the 25.00 mL aliquots.

$n(NaOH)_{titration} = c \times V = 0.105 \times 0.03180 = 3.34 \times 10^{-3}$ mol

titration reaction: $HCl + NaOH \rightarrow NaCl + H_2O$

$n(HCl)_{titration (25.0 mL)} = \left(\frac{1}{1}\right) \times n(NaOH) = 3.34 \times 10^{-3}$ mol

(c)   Calculate the total number of moles of hydrochloric acid present in the 250.0 mL flask and hence calculate the % purity of the chalk.

$n(HCl)_{total excess} = \left(\frac{250}{25}\right) \times 3.34 \times 10^{-3} = 3.34 \times 10^{-2}$ mol

$n(HCl)_{original} = c \times V = 2.00 \times 0.0500 = 0.100$

$n(HCl)_{used in reaction with chalk} = 0.100 - 3.34 \times 10^{-2} = 0.0666$ mol

first reaction: $2 HCl + CaCO_3 \rightarrow CaCl_2 + H_2O + CO_2$

$n(CaCO_3) = \left(\frac{1}{2}\right) \times n(HCl) = \left(\frac{1}{2}\right) \times 0.0666 = 0.0333$ mol

$m(CaCO_3) = n \times M = 0.0333 \times 100.09 = 3.33$ g

$\% \text{ purity} = \left(\frac{3.33}{3.63}\right) \times 100 = 91.7\%$

Acid and Base Titrations Answers

1.   The mixing of accurate volumes of solutions to the point where complete reaction has occurred.

2.   A pipette is used to measure an accurate but fixed volume of liquid. A burette is used to measure accurate volumes of liquid, often in the range of 0 to 50 mL.

3.   A primary standard should be;
   - pure and have a known formula
   - stable to chemical change
   - not change mass during weighing by absorbing or losing water or reacting with air
   - soluble in water (for use with aqueous solutions)
   - have a relatively large formula or molecular mass to reduce errors during weighing.

   Suitable primary standards for use with acid-base titration are anhydrous sodium carbonate and oxalic acid dihydrate.

4.   Absorbs water rapidly from the air during weighing making it impossible to obtain an accurate mass.

5.   Indicators change colour at the end point of a titration. If chosen correctly the indicator's colour change indicates when equivalent amounts of substances have been added and the titration stopped.

6. Equivalence point is when chemically equivalent amounts of the two reactants have been added to the reaction flask. End point is when the indicator changes colour and the titration should be stopped.

7. **(i)**

Sodium carbonate and Hydrochloric acid

(ii)

Sodium Hydroxide and Hydrochloric acid

(iii)

Sodium hydroxide and Ethanoic acid
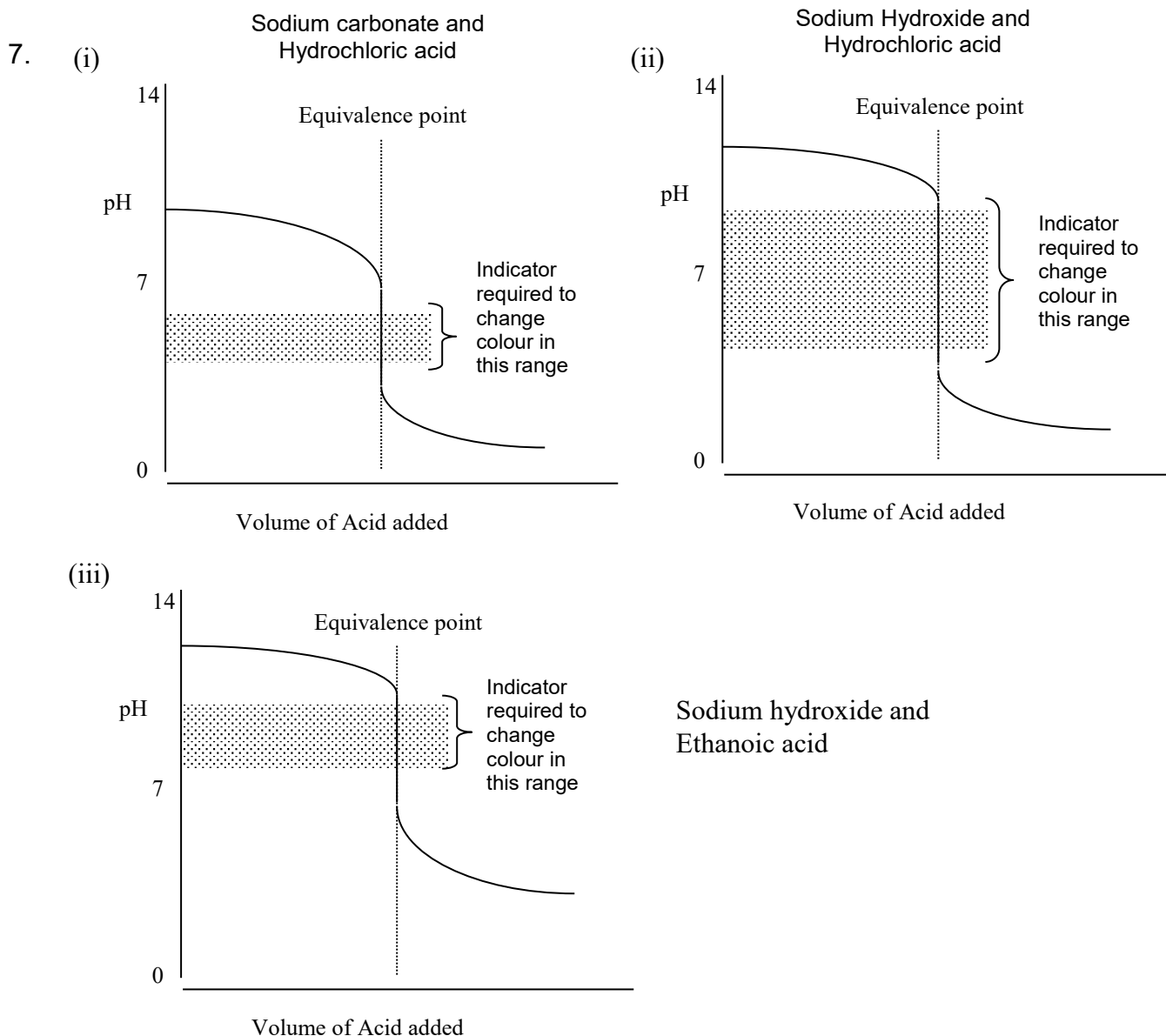
8.   **(i)**   Sodium carbonate and hydrochloric acid.

The pH at equivalence point is acidic as one product hydrolyses (reacts with water).

$$CO_{2(aq)} + H_2O_{(\ell)} \rightleftharpoons HCO_3^-{}_{(aq)} + H^+{}_{(aq)}$$
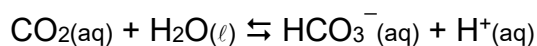
(ii)   hydrochloric acid and sodium hydroxide:

The pH at equivalence point is neutral as neither product hydrolyses.

(iii)   Ethanoic acid and sodium hydroxide

The pH at equivalence point is basic as one product hydrolyses

$$CH_3COO^-{}_{(aq)} + H_2O_{(\ell)} \leftrightarrows CH_3COOH_{(aq)} + OH^-{}_{(aq)}$$
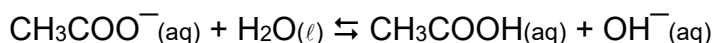
9.   (i)   Methyl orange: changes colour in the range pH = 3 – 4.

      (ii)   Litmus: changes colour in the range pH = 5 – 8
            or Bromothymol Blue: changes colour in the range pH = 6 – 7.6

       Methyl orange (pH 3 – 4) and phenolphthalein (pH 8.3 – 10.0) can also be used as the pH at the equivalence point changes extremely rapidly from a pH of about 8 to 3.

      (iii)   phenolphthalein: changes colour in the range pH 8.3 – 10.0.

10.   In weak acid – weak base titrations both products hydrolyse. The degree of hydrolysis of each product depends on how weak each substance is. If both acid and base are equally weak then the degree of hydrolysis will be the same and the resulting solution will be neutral at equivalence point.

      If the acid is weaker than the base, then the solution at equivalence point will be basic but if the base is weaker the solution will be acidic. Choice of an indicator is therefore difficult.

Acid – Base Neutralisation Answers

1. (i) $n(NaOH) = n(H^+) = n(HC\ell) = 2.00$ mol
    (ii) $n(NaOH) = n(H^+) = 2n(H_2SO_4) = 2 \times 1.00 = 2.00$ mol
    (iii) $n(NaOH) = n(H^+) = 3n(H_3PO_4) = 3 \times 0.500 = 1.50$ mol
    (iv) $n(NaOH) = n(H^+) = n(HNO_3) = 1.40$ mol
    (v) $n(NaOH) = n(H^+) = 2n(H_2SO_4) = 2(0.200) = 0.400$ mol

2. $OH^-_{(aq)} + H^+_{(aq)} \rightarrow H_2O_{(\ell)}$

    $n(H^+) = n(HC\ell) = cV = 2.00 \times 0.0300 = 0.0600$ mol
    $n(NaOH) = n(OH^-) = n(H^+) = 0.0600$ mol
    $$V(NaOH) = \frac{n}{c} = \frac{0.0600}{3.00} = 0.0200 \text{ L} = 20.0 \text{ mL}$$

3. (i)
    $$n(NaOH) = \frac{m}{M} = \frac{2.00}{39.998} = 0.0500 \text{ mol} \qquad [M(NaOH) = 39.998 \text{ g mol}^{-1}]$$
    $$[NaOH] = \frac{n}{V} = \frac{0.0500}{0.250} = 0.200 \text{ mol L}^{-1}$$

    (ii) $m(NaOH)_{\text{in 1.00 L}} = nM = 0.200 \times 39.998 = 8.00$ g L^{-1}

    (iii) $OH^-_{(aq)} + H^+_{(aq)} \rightarrow H_2O_{(\ell)}$

    $n(OH^-) = n(NaOH) = cV = 0.200 \times 0.0250 = 5.00 \times 10^{-3}$ mol
    $$n(H_2SO_4) = \tfrac{1}{2}x(H^+) = \tfrac{1}{2}x(OH^-) = \tfrac{1}{2}x(5.00 \times 10^{-3}) = 2.50 \times 10^{-3} \, mol$$
    $$V(H_2SO_4) = \frac{n}{c} = \frac{2.50 \times 10^{-3}}{0.100} = 0.0250 \, L = 25.0 \, mL$$

    (iv) $m(H_2SO_4) = nM \qquad = 2.50 \times 10^{-3} \times 98.076 \qquad [M(H_2SO_4) = 98.076$ g mol^{-1}]

    $= 0.245$ g
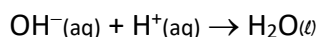
4. (i) $OH^-_{(aq)} + H^+_{(aq)} \rightarrow H_2O_{(\ell)}$
    $n(H^+) = n(HNO_3) = cV = 1.00 \times 0.0200 = 0.0200$ mol
    $n(NaOH) = n(OH^-) = n(H^+) = 0.0200$ mol
    $$V(NaOH) = \frac{n}{c} = \frac{0.0200}{0.400} = 0.0500 \text{ L} = 50.0 \text{ mL}$$

    (ii) $n(H^+) = 2 \times n(H_2SO_4) = 2 \times cV = 2(0.200 \times 0.015) = 6.00 \times 10^{-3}$ mol
    $n(NaOH) = n(OH^-) = n(H^+) = 6.00 \times 10^{-3}$ mol
    $$V(NaOH) = \frac{n}{c} = \frac{6.00 \times 10^{-3}}{0.400} = 0.015 \text{ L} = 15.0 \text{ mL}$$

    (iii) $n(H^+) = 2 \times n(H_2C_2O_4) = 2 \times cV = 2(6.00 \times 10^{-2} \times 0.050) = 6.00 \times 10^{-1}$ mol
    $n(NaOH) = n(OH^-) = n(H^+) = 6.00 \times 10^{-3}$ mol

$$V(\text{NaOH}) = \frac{n}{C} = \frac{6.00 \times 10^{-3}}{0.400} = 0.015 \text{ L} = 15.0 \text{ mL}$$

5. $n(\text{HC}\ell)_{\text{conc}} = cV = 5.00 \times 0.0500 = 0.250 \text{ mol}$

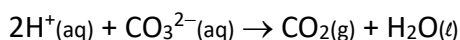$$[\text{HC}\ell]_{\text{Dilute}} = \frac{n}{V} = \frac{0.250}{0.250} = 1.00 \text{ mol L}^{-1}$$

$\text{OH}^- + \text{H}^+ \rightarrow \text{H}_2\text{O}$
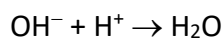
$n(\text{H}^+) = n(\text{HC}\ell) = 1.00 \times 0.0400 = 0.0400 \text{ mol}$

$n(\text{NaOH}) = n(\text{OH}^-) = n(\text{H}^+) = 0.0400 \text{ mol}$

$$V(\text{NaOH}) = \frac{n}{c} = \frac{0.0400}{0.400} = 0.100 \text{ L} = 100 \text{ mL}$$

6.

$$n(\text{H}_2\text{C}_2\text{O}_4.2\text{H}_2\text{O}) = \frac{m}{M} = \frac{12.6}{126.068} \qquad [M(\text{H}_2\text{C}_2\text{O}_4) = 126.068 \text{ g mol}^{-1}]$$
$$= 0.09995 \text{ mol}$$

$$[\text{H}_2\text{C}_2\text{O}_4] = \frac{n}{V} = \frac{0.09995}{4.00} = 0.024987 \text{ mol L}^{-1}$$

$n(\text{H}^+) = 2 \times n(\text{H}_2\text{C}_2\text{O}_4) = 2 \times cV = 2(0.024987 \times 0.020) \qquad = 9.99 \times 10^{-4} \text{ mol}$

$n(\text{KOH}) = n(\text{OH}^-) = n(\text{H}^+) = 9.99 \times 10^{-4} \text{ mol}$

$$[\text{KOH}] = \frac{n}{V} = \frac{9.99 \times 10^{-4}}{0.0360} = 0.0278 \text{ mol L}^{-1}$$

$m(\text{KOH})_{\text{in 1.00 L}} = \ nM \qquad = 0.0278 \times 56.108$

$[M(\text{KOH}) = 56.108 \text{ g mol}^{-1}]$

$= 1.56 \text{ g L}^{-1}$

7.

$$m(\text{H}_2\text{SO}_4) = \frac{98.0}{100} \times 20.0 = 19.6 \text{ g}$$

$$n(\text{H}_2\text{SO}_4) = \frac{m}{M} = \frac{19.6}{98.076} = 0.1998 \text{ mol} \qquad [M(\text{H}_2\text{SO}_4) = 98.076 \text{ g mol}^{-1}]$$

$$[\text{H}_2\text{SO}_4]_{\text{Dilute}} = \frac{n}{V} = \frac{0.1998}{0.250} = 0.799 \text{ mol L}^{-1}$$

$2\text{H}^+_{\text{(aq)}} + \text{CO}_3^{2-}_{\text{(aq)}} \rightarrow \text{CO}_{2\text{(g)}} + \text{H}_2\text{O}_{(\ell)}$
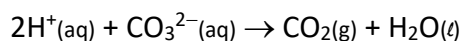
$n(\text{H}^+) = 2 \times n(\text{H}_2\text{SO}_4) = 2 \times cV = 2(0.799 \times 0.010) = 0.01599 \text{ mol}$

$n(\text{Na}_2\text{CO}_3) = n(\text{CO}_3^{2-}) = \tfrac{1}{2}n(\text{H}^+) = \tfrac{1}{2}(0.01599) = 7.99 \times 10^{-3} \text{ mol}$

$$V(\text{Na}_2\text{CO}_3) = \frac{n}{c} = \frac{7.99 \times 10^{-3}}{0.110} = 0.0728 \text{ L} = 72.8 \text{ mL}$$

8. $n(\text{Na}_2\text{CO}_3)_{\text{conc}} = cV = 1.00 \times 0.100 = 0.100 \text{ mol}$

$$[Na_2CO_3]_{Dilute} = \frac{n}{V} = \frac{0.100}{0.250} = 0.400 \text{ mol L}^{-1}$$

$$m(HC\ell) = \frac{6.50}{100} \times 4.00 = 0.260 \text{ g}$$

$$n(HC\ell)_{in\ 4\ g} = \frac{m}{M} = \frac{0.260}{36.458} = 0.00713 \text{ mol} \qquad [M(HC\ell) = 36.458 \text{ g mol}^{-1}]$$

$$V(HC\ell) = 4.00 \times 0.976 = 3.904 \text{ mL}$$

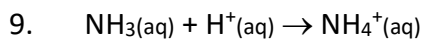$$[HC\ell] = \frac{n}{V} = \frac{0.00713}{3.904 \times 10^{-3}} = 1.826 \text{ mol L}^{-1}$$

$2H^+_{(aq)} + CO_3^{2-}_{(aq)} \rightarrow CO_{2(g)} + H_2O_{(\ell)}$

$n(CO_3^{2-})_{original} = n(Na_2CO_3) = cV = 0.400 \times 0.0200 = 8.00 \times 10^{-3}$ mol
$n(CO_3^{2-})_{used} = \frac{1}{2} \times n(H^+) = \frac{1}{2} \times n(HC\ell)_{in\ 4\ g} = \frac{1}{2}(0.00713) = 3.565 \times 10^{-3}$ mol
$n(CO_3^{2-})_{left} = n(CO_3^{2-})_{original} - n(CO_3^{2-})_{used} = 8.00 \times 10^{-3} - 3.565 \times 10^{-3} = 4.435 \times 10^{-3}$ mol
$n(HC\ell)_{required} = n(H^+) = 2 \times n(CO_3^{2-}) = 2 \times (4.435 \times 10^{-3}) = 8.87 \times 10^{-3}$ mol

$$V(HC\ell) = \frac{n}{c} = \frac{8.87 \times 10^{-3}}{1.826} = 4.85 \times 10^{-3} \text{ L} = 4.85 \text{ mL}$$

9. $NH_{3(aq)} + H^+_{(aq)} \rightarrow NH_4^+_{(aq)}$
      $n(H^+) = n(HNO_3) = cV = 0.200 \times 0.0600 = 0.0120$ mol
      $n(NH_3) = n(H^+) = 0.0120$ mol

$$V(NH_3) = \frac{n}{c} = \frac{0.0120}{0.125} = 0.0960 \text{ L} = 96.0 \text{ mL}$$

10. $Mg_{(s)} + 2H^+_{(aq)} \rightarrow H_{2(g)} + Mg^{2+}_{(aq)}$
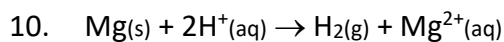
$$n(Mg) = \frac{m}{M} = \frac{1.00}{24.31} = 0.04114 \ mol$$

$$n(H_2SO_4) = \tfrac{1}{2} \times n(H^+) = n(Mg) = 0.04114 \ mol$$

$$[H_2SO_4] = \frac{n}{V} = \frac{0.04114}{0.0250} = 1.65 \ mol \ L^{-1}$$

11. $H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(\ell)}$
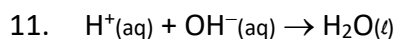
$n(OH^-) = n(NaOH) = cV = 5.00 \times 10^{-2} \times 0.040 = 2.00 \times 10^{-3}$ mol
$n(HNO_3)_{in\ 25\ mL} = n(H^+) = n(OH^-) = 2.00 \times 10^{-3}$ mol

$$n(HNO_3)_{in\ 10.0\ mL\ Dil} = n(HNO_3)_{in\ 250\ mL\ Dil} = \frac{250}{25} \times 2.00 \times 10^{-3} = 2.00 \times 10^{-3} \ mol$$

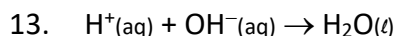$$[HNO_3]_{Conc} = \frac{n}{V} = \frac{2.00 \times 10^{-2}}{0.010} = 2.00 \text{ mol L}^{-1}$$

12. $H^+(aq) + OH^-(aq) \rightarrow H_2O_{(l)}$
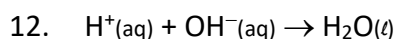
    $n(OH^-) = n(NaOH) = cV = 0.100 \times 0.030 = 0.0030$ mol

    $n(Acid)_{in\ 20\ mL} = n(H^+) = n(OH^-) = 0.0030$ mol

    $n(Acid)_{in\ 3.00\ g} = n(Acid)_{in\ 250\ mL\ dil} = \dfrac{250}{20} \times 0.0030 = 0.0375$ mol

    $M(Acid) = \dfrac{m}{n} = \dfrac{3.00}{0.375} = 80.0$ g mol^{-1}

    Formula mass of acid is 80.0

13. $H^+(aq) + OH^-(aq) \rightarrow H_2O_{(l)}$

    $n(OH^-) = n(NaOH) = cV = 0.100 \times 0.050 = 0.0050$ mol

    $n(Acid)_{in\ 40\ mL} = \frac{1}{3}xn(H^+) = \frac{1}{3}xn(OH^-) = \frac{1}{3}(0.0050) = 1.667 \times 10^{-3}\ mol$

    $n(Acid)_{in\ 2.00\ g} = n(Acid)_{in\ 250\ mL} = \dfrac{250}{40} \times 1.667 \times 10^{-3} = 0.010416\ mol$

    $M(Acid) = \dfrac{m}{n} = \dfrac{2.00}{0.010416} = 192\ g\ mol^{-1}$

    Formula of acid is 192

14. $n(H^+) = 2\ x\ n(H_2SO_4) = 2 \times cV = 2(0.100 \times 0.100) = 0.0200$ mol

    $n(OH^-) = n(NaOH) = cV = 0.500 \times 0.200 = 0.100$ mol

    $n(OH^-)_{left} = 0.100 - 0.0200 = 0.0800$ mol

    $n(H_2SO_4)_{required\ for\ neutralisation} = \frac{1}{2}xn(H^+) = \frac{1}{2}xn(OH^-)_{left} = \frac{1}{2}(0.0800) = 0.0400\ mol$

    $[H_2SO_4]_{used\ to\ complete\ neutralisation} = \dfrac{n}{V} = \dfrac{0.0400}{0.150} = 0.267\ mol\ L^{-1}$

15. $n(H^+) = 2\ x\ n(H_2SO_4) = 2 \times cV = 2(0.200 \times 0.0500) = 0.0200$ mol
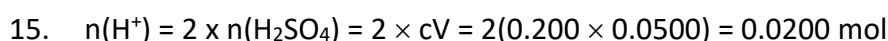
    $n(OH^-) = n(NaOH) = cV = 0.250 \times 0.100 = 0.0250$ mol

    $n(OH^-)_{left} = 0.0250 - 0.0200 = 0.0050$ mol

    $n(H_2SO_4)_{required\ for\ neutralisation} = \frac{1}{2}xn(H^+) = \frac{1}{2}xn(OH^-)_{left} = \frac{1}{2}(0.0050) = 0.00250\ mol$

    $[H_2SO_4]_{used\ to\ complete\ neutralisation} = \dfrac{n}{V} = \dfrac{0.00250}{0.075} = 0.0333\ mol\ L^{-1}$

16. $2H^+(aq) + CO_3^{2-}(aq) \rightarrow CO_{2(g)} + H_2O_{(l)}$
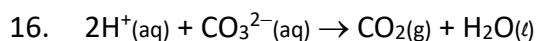
    $n(H^+) = 2\ x\ n(H_2SO_4) = 2 \times cV = 2(6.30 \times 10^{-2} \times 0.025) = 3.15 \times 10^{-3}$ mol

    $n(Na_2CO_3)_{in\ 50\ mL} = n(CO_3^{2-}) = \frac{1}{2}n(H^+) = \frac{1}{2}(3.15 \times 10^{-3}) = 1.57 \times 10^{-3}$ mol

    $n(Na_2CO_3)_{in\ 2.75\ g} = n(Na_2CO_3)_{in\ 300\ mL} = \dfrac{300}{50} \times 1.57 \times 10^{-3} = 9.45 \times 10^{-3}$ mol

    $M(Na_2CO_3.XH_2O) = \dfrac{m}{n} = \dfrac{2.75}{9.45 \times 10^{-3}} = 291$ g mol^{-1}

    $M(Na_2CO_3) = 2(22.99) + 12.01 + 3(16.00) = 105.99 \approx 106$ g mol^{-1}

    $M(XH_2O) = M(Na_2CO_3.XH_2O) - M(Na_2CO_3)$ $= 291 - 106 = 185$ g mol^{-1}

    Number of Water moles $= \dfrac{185}{18.016} = 10.3 \approx 10$

Value of X is 10

17. On heating $NaHCO_3$ decomposes to Na_2CO_3

ie. $2NaHCO_{3(s)} \rightarrow Na_2CO_{3(s)} + CO_{2(g)} + H_2O_{(g)}$

Neutralisation: $2H^+_{(aq)} + CO_3^{2-}_{(aq)} \rightarrow CO_{2(g)} + H_2O_{(l)}$
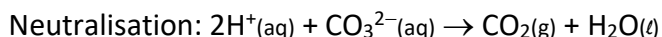
$n(H^+) = 2 \times n(H_2SO_4 = 2 \times cV = 2(0.100 \times 0.01125) = 2.25 \times 10^{-3}$ mol

$n(Na_2CO_3)_{in\ 25.0\ mL} = n(CO_3^{2-}) = \frac{1}{2} \times (H^+) = \frac{1}{2} \times 2.25 \times 10^{-3} = 1.125 \times 10^{-3}$ mol

$n(Na_2CO_3)_{in\ 1.00\ L} = \frac{1000}{25} \times 1.125 \times 10^{-3} = 0.045$ mol

m(NaHCO₃) = nM = 0.900 × 84.008   [M(NaHCO₃) = 84.008 g mol⁻¹]

= 7.56 g

18. $H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O$

n(H⁺) = n(HCℓ) = cV = 0.450 × 0.0300 = 0.0135 mol

n(NaOH)ₑₓcₑₛₛ = n(OH⁻) ₑₓcₑₛₛ = n(H⁺) = 0.0135 mol

$NH_4^+_{(aq)} + OH^-_{(aq)} \leftrightarrows NH_{3(g)} + H_2O_{(l)}$
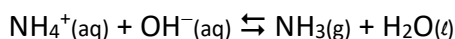
n(OH⁻)ₐddₑd to ammonium salt + n(NaOH) ₐddₑd to ammonium salt = cV = 1.10 × 0.0250 = 0.0275 mol

n(NH₃) = n(OH⁻)ᵣₑₐcₜₑd = n(OH⁻) ₐddₑd to ammonium salt − n(OH⁻)ₑₓcₑₛₛ

=

0.0275 − 0.0135 = 0.0140 mol

m(NH₃) = nM = 0.014 × 17.034

[M(NH₃) = 17.034 g mol⁻¹]

= 0.2385 g

$\%NH_3 = \frac{0.2385}{0.700} \times 100 = 34.1\%$

19. In the first experiment: $H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O$

n(OH⁻) = n(NaOH) = cV = 2.00 × 0.020 = 0.040 mol

n(HCℓ) = n(H⁺) = n(OH⁻) = 0.040 mol

$[HC\ell] = \frac{n}{V} = \frac{0.040}{0.0100} = 4.00$ mol L⁻¹

In second experiment: $2H^+_{(aq)} + CO_3^{2-}_{(aq)} \rightarrow CO_{2(g)} + H_2O_{(l)}$
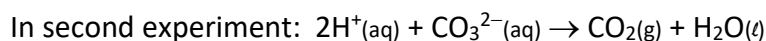
n(H⁺)ₐddₑd to calcite = n(HCℓ) ₐddₑd to calcite = cV = 4.00 × 0.01 = 0.04 mol

n(H⁺)ₑₓcₑₛₛ = n(OH⁻)ᵣₑₐcₜₑd with excess = n(NaOH)ᵣₑₐcₜₑd with excess = cV   = 2.00 × 0.01025

=

0.0205 mol

n(H⁺) ᵣₑₐcₜₑd with Calcium carbonate = n(H⁺)ₐddₑd to calcite − n(H⁺)ₑₓcₑₛₛ

= 0.0400 − 0.0205 = 0.0195 mol

$$n(CaCO_3)_{in\ calcite} = n(CO_3^-) = \tfrac{1}{2}xn(H^+)_{reacted\ with\ CaCO_3} = \tfrac{1}{2}(0.0195) = 0.00975\ mol$$

m(CaCO₃)$_{in\ calcite}$ = nM = 0.00975 × 100.09

[M(CaCO₃) = 100.09 g mol⁻¹]

= 0.9759 g

$$\%CaCO_3 = \frac{0.9259}{1.00} \times 100 = 97.6\%$$

20. Zn$_{(s)}$ + 2H⁺$_{(aq)}$ → H₂$_{(g)}$ + Zn²⁺$_{(aq)}$
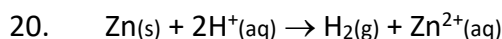
n(H⁺) = n(HCℓ) = cV = 0.250 × 0.0380 = 9.50 × 10⁻³ mol

n(Zn) = ½ x n(H⁺) = ½ (9.50 × 10⁻³) = 4.75 × 10⁻³ mol

m(Zn) = nM = 4.75 × 10⁻³ × 65.38 = 0.311 g

$$\%Zn = \frac{0.311}{0.325} \times 100 = 95.6\%$$

21. 2H⁺$_{(aq)}$ + CO₃²⁻$_{(aq)}$ → CO₂$_{(g)}$ + H₂O$_{(ℓ)}$
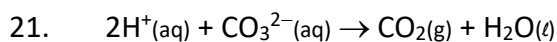
$$n(CO_3^{2-}) = n(Na_2CO_3) = \frac{m}{M} \qquad\qquad [M(Na_2CO_3) = 105.99\,g\,mol^{-1}]$$

$$= \frac{0.200}{105.99} = 1.887 \times 10^{-3}\,mol$$

$$n(H_2SO_4) = \tfrac{1}{2}xn(H^+) = \frac{1}{2} \times 2xn(CO_3^{2-}) = \tfrac{1}{2} \times 2 \times 1.887 \times 10^{-3} = 1.887 \times 10^{-3}\,mol$$

$$[H_2SO_4] = \frac{n}{V} = \frac{1.887 \times 10^{-3}}{0.0392} = 0.0481\,mol\ L^{-1}$$

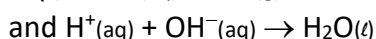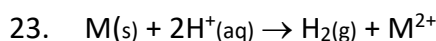22. 2Na$_{(s)}$ + 2H₂O$_{(aq)}$ → H₂$_{(g)}$ + 2Na⁺$_{(aq)}$ + 2OH⁻$_{(aq)}$

and H⁺$_{(aq)}$ + OH⁻$_{(aq)}$ → H₂O$_{(aq)}$
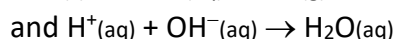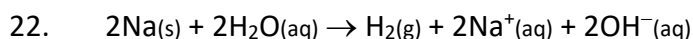
$$n(Na) = \frac{m}{M} = \frac{0.0920}{22.99} = 4.002 \times 10^{-3}\,mol$$

$$n(HC\ell) = n(H^+) = n(OH^-) = n(Na) = 4.002 \times 10^{-3}\ mol$$

$$V(HC\ell) = \frac{n}{c} = \frac{4.002 \times 10^{-3}}{0.250} = 0.0160\ L = 16.0\ mL$$

Suitable indicators would be litmus, bromothymol blue, or even phenolphthalein and methyl orange as both acid and base are strong.
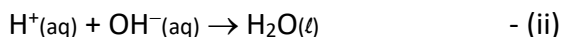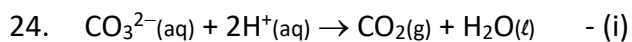
23. M$_{(s)}$ + 2H⁺$_{(aq)}$ → H₂$_{(g)}$ + M²⁺

and H⁺$_{(aq)}$ + OH⁻$_{(aq)}$ → H₂O$_{(ℓ)}$

n(H⁺)$_{added\ to\ metal}$ = n(HCl) $_{added\ to\ metal}$ = cV = 0.500 × 0.100 = 0.0500 mol

n(H⁺)$_{excess}$ = n(OH⁻) = n(NaOH) = cV = 0.250 × 0.0680 = 0.0170 mol

n(H⁺)$_{reacted\ with\ metal}$ = n(H⁺)$_{added\ to\ metal}$ − n(H⁺)$_{excess}$

= 0.0500 − 0.0170 = 0.0330 mol

$$n(M) = \tfrac{1}{2}n(H^+)_{\text{reacted with metal}} = \tfrac{1}{2}(0.0330) = 0.0165 \text{mol}$$

$$M = \frac{m}{n} = \frac{0.400}{0.0165} = 24.2$$

24. CO_3^{2-}(aq) + $2H^+$(aq) $\rightarrow$ $CO_{2(g)}$ + $H_2O_{(\ell)}$ - (i)
      H^+(aq) + OH^-(aq) $\rightarrow$ $H_2O_{(\ell)}$ - (ii)

$$n(CaCO_3) = \frac{m}{M} = \frac{0.100}{100.09} = 9.991 \times 10^{-4} \text{ mol} \qquad [M(CaCO_3) = 100.09 \text{ g mol}^{-1}]$$

$n(H^+)_{\text{reacted}}$ = 2 x n($CaCO_3$) = 2(9.991×10^{-4}) = 1.9982×10^{-3} mol
$n(H^+)_{\text{added to acid}}$ = n(HCl)$_{\text{added to acid}}$ = cV = [HCℓ] $\times$ 0.0250 mol
$n(H^+)_{\text{excess}}$ = n(OH^-)$_{\text{reacted with excess HCl}}$ = n(NaOH)$_{\text{reacted with excess acid}}$ = cV
$\qquad\qquad\qquad\qquad\qquad$ = [NaOH] $\times$ 0.00550
$n(H^+)_{\text{excess}}$ also = $n(H^+)_{\text{added to acid}}$ $-$ $n(H^+)_{\text{reacted}}$
$\qquad\qquad$ = [HCℓ] $\times$ 0.0250 $-$ 1.9982×10^{-3}

so [NaOH] $\times$ 0.00550 = [HCℓ] $\times$ 0.0250 $-$ 1.9982×10^{-3} (equation i)

From 2$^{\text{nd}}$ Titration n(OH^-) = n(NaOH) = cV = [NaOH] $\times$ 0.0275
And n(H^+) = n (HCℓ) = cV = [HCℓ] $\times$ 0.0250

n(OH^-) = n(H^+)
$\therefore$ [NaOH] $\times$ 0.0275 = [HCℓ] $\times$ 0.0250

$$ie \ [NaOH] = \frac{[HCl] \ x \ 0.0250}{0.0275} \qquad\qquad\qquad \text{(equation   ii)}$$

Substituting equ (ii) into equ (i) and solving for [HCℓ]

$$\frac{[HCl] \ x \ 0.0250}{0.0275} \ x \ 0.00550 = [HCl]x \ 0.0250 - \ 1.9982 \ x \ 10^{-3}$$

$$[HCl] = 0.0999 \text{ mol L}^{-1} \ \text{ and } [NaOH] = 0.908 \text{ mol L}^{-1}$$